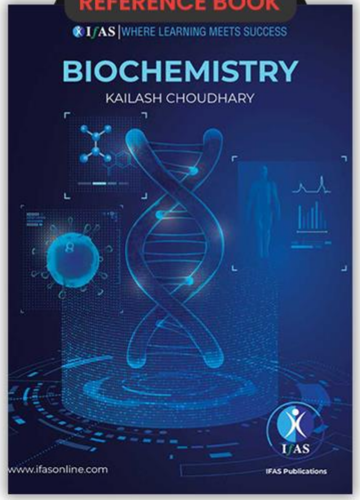
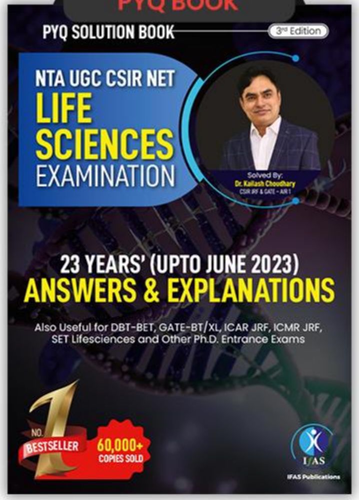
"नामुमकिन कुछ भी नही है हम वह सब कर सकते हैं जो हम सोच सकते हैं और वह सब सोच सकते हैं जो हमने कभी नहीं सोचा।"

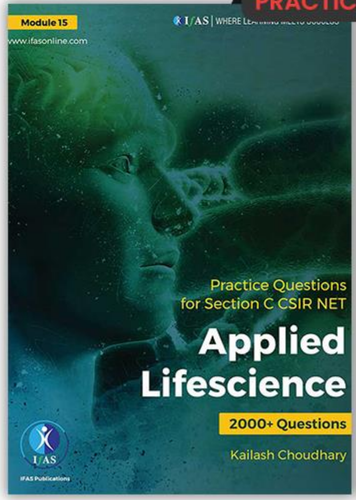
CSIR NET – Life Science

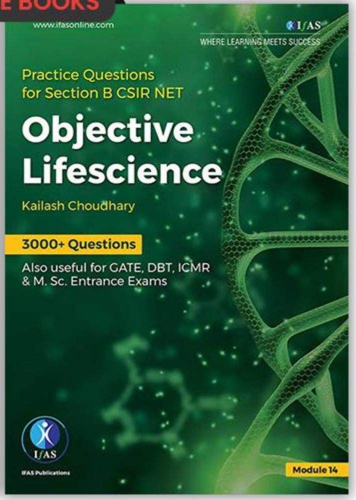
Unit 1: Biochemistry

04 pH and Buffer









Points to be covered in this Lecture

- What are acid and Bases? ✓
- Ionic Product of Water ✓
- What is pH? ✓
- pH of Acid and Base ✓
- Ka and pKa ✓
- Buffers ✓

Basics of log calculations

Number	Logarithm
1	0
2	0.3
3	0.48
4	0.6
5	0.7
6	0.78
7	0.86
8	0.9
9	0.95
10	1.0

Number	Logarithm
10^0	0
10^1	1
10^2	2
10^3	3
$1/10 \rightarrow 10^{-1}$	-1
$1/100 \rightarrow 10^{-2}$	-2
10^{-3}	-3
$10^{2.2}$	2.2
$10^{-2..5}$	-2.5
10^{-6}	-6

$\log 1 = 0$
$\log 10 = 1$
$\log 100 = 2$
$\log 1000 = 3$

$\log \frac{1}{10} = -1$

$\log \frac{1}{100} = -2$

Calculate

Log of 50

$\log 5 \times 10$

$\log 5 + \log 10$

$= 0.7 + 1$

$= 1.7$

Log 1/20

$\log 1 - \log 20$

$0 - \log 20$

$-(\log 20)$

$-(\log 2 \times 10)$

$-(\log 2 + \log 10)$

$-(0.3 + 1)$

$-(1.3)$

-1.3

Log 0.2

$\log \frac{2}{10}$

$\log 2 - \log 10$

$= 0.3 - 1$

$= -0.7$

$\log \frac{1}{5}$

$= -0.7$

Calculate

Log 5.05×10^{-5}

$\log 5 \times 10^{-5}$

$\log 5 + \log 10^{-5}$

$= 0.7 - 5$

$= -4.3$

Log 4.05×10^{-6}

$\log 4 \times 10^{-6}$

$\log 4 + \log 10^{-6}$

$= 0.6 - 6$

$= -5.4$

Antilog

What is antilog of 2? $\longrightarrow$ 100

What is antilog of -1 ? $\longrightarrow$ $\frac{1}{10}$ or 10^{-1}

What is antilog of 0.6 ? $\longrightarrow$ 4

What is antilog of -0.7 ? $\longrightarrow$ $\frac{1}{5}$ or 0.2 or 2×10^{-1}

What is antilog of 1.3 ?

Antilog 1 + Antilog 0.3

$10 \times 2 = 20$

Antilog 1.6

$10 \times 4 = 40$

Antilog -2.7

$= 1/100 \times 5$

$= 1/500$

What is acid and Base?

Any hydrogen-containing substance that is capable of donating a proton (hydrogen ion) to another substance. or water

$HA \rightarrow H^+ + A^-$

Acid Base

A base is a molecule or ion able to accept a hydrogen ion from an acid.

$A^- + H^+ \rightarrow HA$

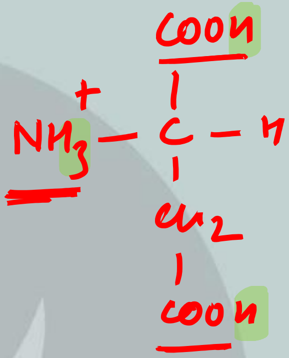
Base Acid

- Acid donate H^+
- Acid accept electron
- Acid pH is less than 7
- Acid is sour in taste

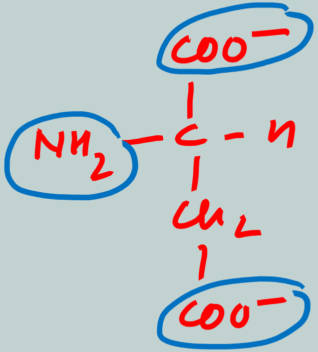
- Accepts H^+
- donate electron
- pH is higher than 7

Example of Acid and its Conjugate Base

Acid	Conjugate Base
HCl	Cl-
Acetic Acid (CH_3COOH)	Acetate (CH_3COO^-)
CO_2 (H_2CO_3)	Bicarbonate (HCO_3^-)
Citric Acid	Citrate
Formic Acid	Formate
Lactic Acid	Lactate
Ammonium Ion (NH_4^+)	Ammonia (NH_3)



3 Acidic group



3 basic group

Ionic product of water (Kw)

$H_2O \longrightarrow \underline{55.56\ M}$

The product of concentration of H^+ ion and OH^- ions.

$$\boxed{K_w = [H^+][OH^-]}$$
$$= (1\times10^{-7})(1\times10^{-7})$$
$$K_w = 1\times10^{-14}$$

25°C

$$H_2O \xrightarrow[\text{Temp}^r]{25^\circ C} H^+ + OH^-$$

55.56 M → 10^{-7}, 10^{-7} M

0.0000001

Temp^r ↑ : more Dissociation

Temp^r ↓ : lesser Dissociation

If concentration of H^+ is given concentration of OH^- can be calculated and vice-versa

$K_w = 10^{-14}\ M^2$

$K_w = (H^+) \times (OH^-)$

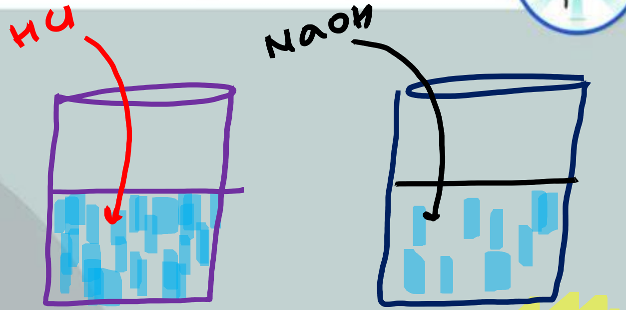
$10^{-14} = (H^+) \times (OH^-)$

$(H^+) = \frac{10^{-14}}{(OH^-)}$

$(OH^-) = \frac{10^{-14}}{[H^+]}$

Acid or Base in aqueous environment

- Only one acid in water of H_3O^+ ion
- Only one base in water OH^- ions



$$HA \longrightarrow H^+ + A^-$$

$$H^+ + H_2O \rightleftharpoons H_3O^+$$

$$HCl + H_2O \longrightarrow Cl^- + H_3O^+$$

$$NaOH + H_2O \longrightarrow Na^+ + OH^-$$

$$H_3O^+ \rightleftharpoons H_2O + H^+$$

$$H_2O \rightleftharpoons OH^- + H^+$$

Strong acid and Weak Acids

A strong acid fully dissociates (100 %) into its ions in water.
Eg. HCl, HI

$$HA \xrightarrow{100\%} H^+ + A^-$$

A weak acid partially dissociated into its ions in an aqueous solution or water.

$$HA \xrightarrow[0.1\%]{1\% \quad 10\%} H^+ + A^-$$

Eg. Acetic Acid
Formic Acid
Citric Acid

What is pH?

- Power of hydrogen in aqous solution
- Inverse relationship with H^+ concentration
- $pH = -\log[H^+]$

$H^+ + H_2O \rightarrow \boxed{H_3O^+}$

$pH = -\log[H_3O^+]$

Molar = $\frac{\text{mole}}{\text{lt}}$

$pH \propto \frac{1}{[H^+]}$

$pH = \log\frac{1}{[H^+]}$

$pH = \log 1 - \log(H^+)$

$= 0 - \log(H^+)$

$= -\log(H^+)$

Apply your mind:

Question: A strong acid solution with pH 4 was diluted to 100 times, what will be the pH of diluted solution?

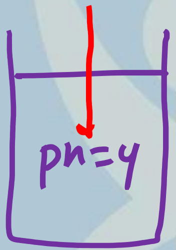
(1) 2

(2) 4

(3) 6

(4) 8

100 times dilution (H_2O)



$H^+ = \downarrow$ pH = $\uparrow$

pH will increase by 2 units

Apply your mind:

The pH of cytosol is 7.2 and the pH of lysosome is 4.2. The cytosol has a $[H^+]$ that is

(1) 10 times lower than that of lysosome

(2) 0.1 times lower than that of lysosome

(3) 100 times lower than that of lysosome

(4) 1000 times lower than that of lysosome.

pH	cytosol	Lysosome
	7.2	4.2
	$H^+ = \downarrow$	$H^+ = \uparrow$

Difference 3 unit

Conc^n = 1000 fold variation

10 fold gradient
100 " "
1000 " "
10,000 " "

pH change
1 unit
2 unit
3 unit
4 unit

Why pH of distilled water kept in beaker under pH meter is not exactly 7?

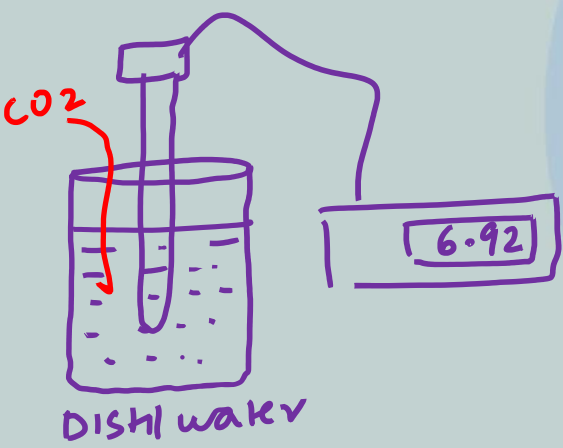
Why the pH of rain water is around 5.6?

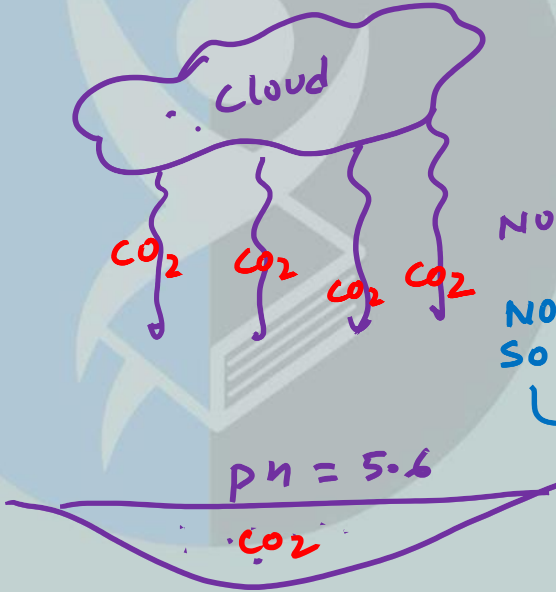
$CO_2 + H_2O \longrightarrow H_2CO_3$

(Acid)

$H_2CO_3 \rightarrow H^+ + HCO_3^-$

H^+ → Decrease pH





No pollution

NOx, SOx] pollutant

→ pH = lower than 5 (Acid Rain)

pH of mono-protonic strong acid

$HCl \xrightarrow{100\%} H^+$, $HBr \xrightarrow{100\%} H^+$, $HI \xrightarrow{100\%} H^+$

Acid concentration directly provide concentration of H^+

$pH = -\log[H^+]$

- 0.1 M HCl ?

$H^+ = 0.1$ or 10^{-1}

$pH = -\log(H^+)$
$= -\log 10^{-1}$
$= -(-1)$
$= 1$

Apply your mind:

Question: Calculate pH of 1mM HCL

$$(H^+) = 1\ mM$$
$$= 1 \times 10^{-3}\ M$$
$$= 10^{-3}\ M$$

$$pH = -\log [H^+]$$
$$= -\log 10^{-3}$$
$$= -(-3)$$
$$= +3$$

pH of multi-protonic strong acid

Concentration of H^+ = Acidity x Acid concentration

2 ×
3 ×

pH= -log[H^+]

$= -\log 2 \times 10^{-3}$
$= -(\log 2 + \log 10^{-3})$
$= -(0.3 - 3)$
$= -(-2.7)$
$= +2.7$

H_2SO_4 H_3PO_4
Acidity = 2 3

• pH of 1 mM H_2SO_4

$H^+ = 2 \times 1\ mM$
$= 2\ mM$
$= 2 \times 10^{-3}\ M$

Apply your mind:

Normality of Acid = H^+ concn

Calculate pH of 0.05 M H_2SO_4?

(1) 0.7

(2) 1

(3) 1.3

(4) 2

$[H^+] = 2 \times$ Acid concn

$= 2 \times 0.05\ M$

$= 0.1\ M = 10^{-1}\ M$

$pH = -\log 10^{-1}$

$= -(-1)$

$= +1$

pH of strong base:

$pOH = -\log[OH^-]$

$pOH = \log \frac{1}{(OH^-)}$

$[OH^-] \longrightarrow pOH$

$pH + \underline{pOH} = \underline{14}$

$pH = 14 - pOH$

$OH^- \longrightarrow H^+$

$\downarrow$

pH

$[H^+] = \frac{10^{-14}}{(OH^-)}$

Apply your mind:

Question: Calculate the pH of 0.1 M KOH

Base

(1) 1
(2) 6.98
(3) 7.02
(4) 13

$OH^- = 0.1 = 10^{-1}$

$pOH = -\log[OH^-]$

$= -\log 10^{-1}$

$pOH = 1$

$pH = 14 - pOH$

$= 14 - 1$

$= 13$

$H^+ = \frac{10^{-14}}{(OH^-)}$

$H^+ = \frac{10^{-14}}{10^{-1}}$

$H^+ = 10^{-13}$

$pH = -\log(H^+)$

$pH = -\log 10^{-13}$

$= -(-13)$

$= 13$

Acid pH is always lesser than 7

Base pH is always higher than 7 at 25°C

Apply your mind:

$H^+ \times OH^- = 10^{-14}$

The (OH^-) of 0.1 N HCl solution will be

$H^+ = 10^{-1}$

$[OH^-] = \frac{10^{-14}}{[H^+]}$

$= \frac{10^{-14}}{10^{-1}}$

$= 10^{-13}$

$pH = 1$

$pOH = 14 - 1$

$pOH = 13$

$[OH^-] = 10^{-13}$

Acid is highly concentrated

pH scale

0 ← 7 → 14

Theoretically pH can be less than zero (negative)

Calculate pH of 10 M HCL solution?

$H^+ = 10\ M$

$pH = -\log 10$

$= -1$

Calculate pH of 100 M HBr solution?

$H^+ = 100$

$pH = -\log(100)$

$= -2$

pH of aqueous solutions ranges from 0-14 at 25^0 C

But pH of pure water can also vary with temperature

$H_2O \xrightarrow{70^{\circ}} H^+ + OH^-$

25°C

At lower temperature pH + pOH can be higher than 14

At higher temperature pH + pOH can be lower than 14

$\downarrow pH \propto \frac{1}{Temp^r} \uparrow$ $\downarrow pOH \propto \frac{1}{Temp^r} \uparrow$

$H^+ = \uparrow$ $pH = \downarrow$

$OH^- = \uparrow$ $pOH = \downarrow$

Sum of pH + pOH will be less than 14

10^{-8} M

pH of Very dilute solution of strong acid

HCl

Total H^+ Concentration will be

- H^+ from dilute acid and H^+ of water (10^{-7} M)
- It would be slightly lesser than 7

10^{-8} M

$(H^+)_{HCl} = 0.00000002$ m (10^{-8})

$(H^+)_{H_2O} = 0.0000002$ m (10^{-7})

$pH = -\log(\text{Total } H^+)$

Apply your mind:

Acid pH must be lesser than 7.

Calculate pH of 10^{-8} M HCl (Given log of 1.1=0.04)

(1) 6.96

(2) 7.04

(3) 7.0

(4) 8.0

$H^+_{HCl} = 10^{-8}$ M

$H^+_{H_2O} = 10^{-7}$ M

pH = − log (H^+)
= − log 10^{-8}
= 8 ← Not possible for acid

Total(H^+) = $10^{-8} + 10^{-7}$
= $(10^{-1} \times 10^{-7}) + (1 \times 10^{-7})$
= $0.1 \times 10^{-7} + 1 \times 10^{-7}$
= $(0.1 + 1) \times 10^{-7}$
= 1.1×10^{-7}

pH = − log 1.1×10^{-7}
= − log 1.1 + log 10^{-7}
= −(0.04 − 7)
= −(−6.96)
= 6.96

Apply your mind:

Base

Calculate the pH of 10^{-8} M NaOH solution?

Low concr

(1) 4 ✗

(2) 6.96 ✗

(3) 7.04 ✓

(4) 10 ✗

$pOH = -\log 10^{-8}$

$= 8$

$pH = 14-8$

$= 6 \rightarrow$ which is not possible for base

must be higher than 7

pH of mixture of two acidic solution when added in equal volume

- pH depends on H^+ ion from solution 1 + solution 2
- It will in between pH of solution 1 and solution 2
- pH will be not average
- It will be closer to strong acidic solution (near lower pH)

pH 3 + pH 5
100 ml 100 ml

Answer : 3.3

3 ⟵ ⟶ 5

4

$\frac{3+5}{2} = \frac{8}{2} = 4$

Apply your mind:

What will be of mixture if you add 100 mL of pH=2 HCl solution and 100 mL of pH=4 HCl solution?

(1) ~~2~~
(2) 2.3
(3) ~~3.0~~
(4) ~~3.7~~

- In betn 2 to 4
- It will not be average i.e 3
- It would near to 2

provides more H^+

pH of mixture of two basic solution when added in equal volume

- pH depends on H^+ ion from solution 1 ($pH = 8$) + solution 2 ($pH = 10$)
- It will in between pH of solution 1 and solution 2 — betn 8 to 10
- pH will be not average $\neq 9$
- It will be closer to strong basic solution (higher pH) — near 10 — provide more OH^-

What will be of mixture if you add 100 mL of pH=8 NaOH solution and 100 mL of pH=12 NaOH solution?

(1) 8.0

(2) 8.3

~~(3) 10.0~~

(4) 11.7

pH = 8 pH = 12

• In betⁿ 8 to 12

• will not be average $\frac{12+8}{2} = \frac{20}{2} = 10$

• Near strong base (pH=12)

pH of mixture of acid + base solution when added in Equal Volume

H^+ OH^-

- pH depends on H^+ ion from 4 solution 1 + OH^- ion 10 solution 2
- It will in between pH of solution 1 and solution 2 4 to 10
- pH will be average
- It will be closer to 7

$$\frac{4+10}{2} = \frac{14}{2} = 7$$

Apply your mind:

If 500 ml of pH 3 and 500 ml of pH 11 solution are mixed, then what would be resultant pH of mixture?

(1) 7.0
(2) 3.3
(3) 10.7
(4) 3.0

500 mL pH = 3 Acid
500 mL pH = 11 Base

$$pH = \frac{3+11}{2} = \frac{14}{2} = 7$$

Ka of weak acid

Dissociation constant for acid

$$HA \longrightarrow H^+ + A^-$$

$$Ka = \frac{[H^+][A^-]}{[HA]}$$

$$H^+ = \sqrt{Ka \cdot [HA]}$$

$$pH = -\log(H^+)$$

- High Ka means strong acid
- Low Ka means weak acid

$pKa = -\log Ka$

$pH \propto \frac{1}{H^+}$

$pH = \log \frac{1}{(H^+)}$

$pH = -\log (H^+)$

$pKa \propto \frac{1}{Ka}$

$pKa = \log \frac{1}{Ka}$

$pKa = -\log Ka$

pKa is a **pH** at which the concentration of **weak acid** and its conjugate base will be in equimolar concentrations.

Acid Base Base

$CH_3COOH \longrightarrow CH_3COO^-$

pH = ↑

0.5 M (50% Dissociation) 0.5

- Strong acid has low pKa, Higher Ka
- Weak acid has higher pKa, low Ka

Mono-protonic acid has Single pKa

$$CH_3COOH \xrightleftharpoons{pH = 4.76} CH_3COO^- + H^+$$

50% 50%

Acetic acid → $pKa = 4.76$

- Formic acid
- Lactic acid
- Pyruvic acid

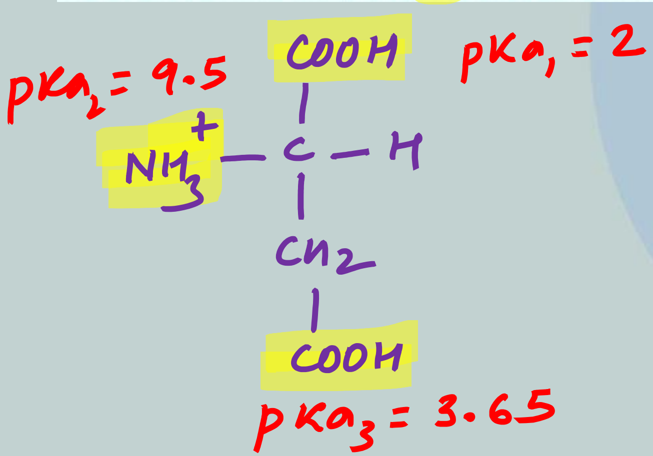
Poly-protonic acid has Multiple pKa

$$H_3PO_4 \underset{}{\overset{pK_a\ 2.1}{\rightleftharpoons}} H_2PO_4^- + H^+ \overset{pK_a\ 6.8}{\rightleftharpoons} HPO_4^{-2} + H^+ \overset{pK_a\ 12.3}{\rightleftharpoons} PO_4^{-3} + H^+$$

$pKa_1 = 2.1$

$pKa_2 = 6.8$

$pKa_3 = 12.3$



$pKa_1 = 6.1$ $\quad$ $pKa_2 = 10.8$

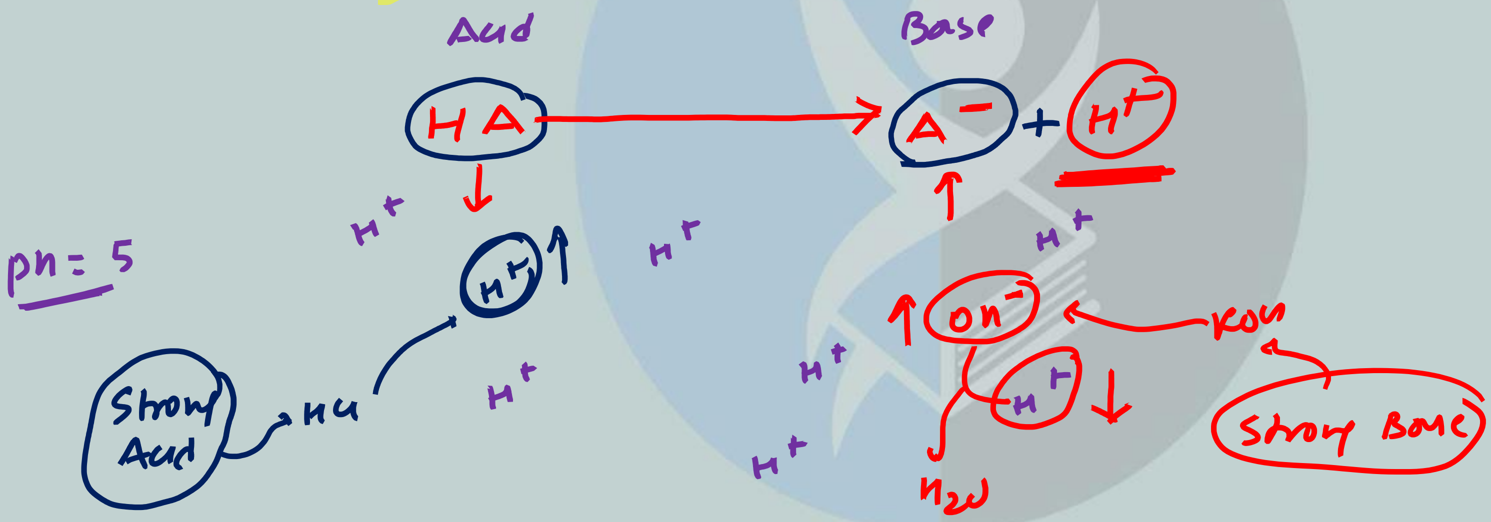
$$H_2CO_3 \rightleftharpoons HCO_3^- \rightleftharpoons CO_3^-$$

Carbonic acid $\quad$ Bicarbonate Ion $\quad$ -Carbonate Ion

What are buffers?

- It is mixture of weak acid and its conjugate base

- Retard the change in the pH of a solution on addition of strong acids or strong bases .



Henderson - Hasselbalch equation :

$HA \longrightarrow A^- + H^+$

HA: Acid Protonated form

A^-: Base Deprotonated form

$$\underline{pH} = \underbrace{pK_a}_{\text{of weak acid}} + \log \frac{[A^-] \leftarrow \text{Base form}}{[HA] \leftarrow \text{Acid form}}$$

- Best Buffering is when pH of buffer is equal to pKa of weak acid
- General Buffering range is $\boxed{pH = pKa \pm 1}$
- When there is more than one pKa choose pKa closest to desired pH

[Acid] = [Base]   log 1 = 0 $\boxed{pH = pKa}$

Acetic acid (pKa = 4.76) → Best buffer pH = 4.76

↳ Buffer Range = 4.76 ± 1 = 3.76 to 5.76

How to make buffer of required pH if pKa of acid is given?
What must ratio of base to acid?

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

100 ← $\frac{\text{Base}}{\text{Acid}}$
1

- [base]>[acid] means pH>pKa
- [base]=[acid] means pH=pKa
- [base]<[acid] means pH<pKa

Base/Acid	pH
10 : 1	pKa + 1
100 : 1	pKa + 2
1000 : 1	pKa + 3
1 : 10	pKa − 1
1 : 100	pKa − 2
1 : 1000	pKa − 3

Apply your mind:

Prepare a 0.2 M acetate buffer with pH 4.76 (pKa=4.76)

(1) 0.1 M Acetate + 0.1 M Acetic acid
(2) 0.1 M Acetate + 0.01 M Acetic acid
(3) 0.01 M Acetate + 0.1 M Acetic acid
(4) 0.1 M Acetate + 0.2 M Acetic acid

Base Acid

pH = pKa

Base = Acid

$$pH = pKa + \log \frac{\text{Acetate}}{\text{Acetic acid}}$$

$$pH - pKa = \log \frac{\text{Acetate}}{\text{Acetic acid}}$$

$$\text{Antilog}\,(pH - pKa) = \frac{\text{Acetate}}{\text{Acetic acid}}$$

$$\text{Antilog}\,(4.76 - 4.76) = \frac{\text{Base}}{\text{Acid}}$$

$$\text{Antilog}\ 0 = \frac{\text{Base}}{\text{Acid}}$$

$$1 = \frac{\text{Base}}{\text{Acid}}$$

Apply your mind:

Prepare Acetate buffer with pH 5.76 (pKa=4.76)

(1) 0.1 M Acetate + 0.1 M Acetic acid ✗
(2) 0.1 M Acetate + 0.01 M Acetic acid ✓
(3) 0.01 M Acetate + 0.1 M Acetic acid ✗
(4) 0.1 M Acetate + 0.2 M Acetic acid ✗

$$\frac{Base}{Acid} = \frac{0.1}{0.01} = \frac{10}{1}$$

$$pH = pKa + \log\frac{B}{A}$$

$$5.76 = 4.76 + \log\frac{B}{A}$$

$$= 4.76 + \log 10$$

$$= 4.76 + 1$$ ✓

$$\text{Antilog}(pH - pKa) = \frac{Base}{Acid}$$

$$\text{Antilog}(5.76 - 4.76) = B/A$$

$$\text{Antilog } 1 = B/A$$

$$\frac{10}{1} = B/A$$

Apply your mind:

100 mL of 0.1 M sodium acetate has pH of 8.9. To this solution 1 mL of 1 M acetic acid (pKa =4.76) of pH 2.80 is added. The pH of resultant mixture will be?

(1) 8.9
(2) 3.76
(3) 4.76
(4) 5.76

Base 100 mL 0.1 M $0.1 \frac{mole}{1000\,mL}$

1000 mL = 0.1
$1\,mL = \frac{0.1}{1000}$

Base → 100 ml $= \frac{0.1}{1000} \times 100 = 0.01 = 10^{-2}$ mole

Acid 1 ml 1 M 1 mole/1000 mL

1000 ml → 1 mole
1 ml → $\frac{1}{1000}$ mole $= 10^{-3}$ mole

$pH = pKa + \log \frac{Acetate}{Acetic\ acid}$

$= 4.76 + \log \frac{10^{-2}}{10^{-3}}$

$= 4.76 + \log 10^{1}$

$= 4.76 + 1 = 5.76$

If pH of buffer is given we can estimate base/acid ratio

$$pH = pK_a + log\frac{[A^-]}{[HA]}$$

Antilog (pH-pKa) = $\frac{Base}{Acid}$

Calculate [base]/[acid] ratio using this equation.

$$\frac{[Base]}{[Acid]} = \frac{[A^-]}{[HA]} = Antilog\ (pH - pK_a)$$

Apply your mind

The apparent pH of a fluid is 7.45, where bicarbonate buffer is involved for maintaining its pH. Values of pKa of carbonic acid are 6.15 and 10.45. The molar ratio of [conjugate base]:[acid) is (Hint: antilog 1.3 = 20.0, and antilog 10^3 = 1000)

(1) 1: 20
(2) 20: 1
(3) 1: 1000
(4) 1000: 1

$pKa_1 = 6.15$

$pKa_2 = 10.45$ ×

$pH = 7.45$ $pKa = 6.15$

$\text{Antilog}(pH - pKa) = \frac{\text{Base}}{\text{Acid}} = \frac{HCO_3^-}{CO_2} = \frac{20}{1}$

$\text{Antilog}(7.45 - 6.15) = B/A$

$\text{Antilog } 1.3 = B/A$

$20/1 = B/A$

Total = P + DP

11 = 10 + 1

Degree of protonation

$$pH = pK_a + log\frac{[Deprotnated\ form]}{[Protonated\ form]}$$

$\leftarrow A^-$

$\leftarrow HA$

$\frac{1}{10}$

$P = \frac{10}{11} \times 100 = 90\%$

$DP = \frac{1}{11} \times 100 = 10\%$

pH and pKa	Base/Acid Ratio	Deprotonated Form	Protonated Form
pH=pKa	1:1	50%	50%
pH= pKa -1	1:10	10%	90%
pH=pka-2	1:100	1%	99%
pH=pka+1	10:1	90%	10%
pH=pka+2	100:1	99%	1%

lower pH

Higher pH

Apply your mind

100 ml of 0.02 M acetic acid (pK_a= 4.76) is titrated with 0.02 N KOH. After adding some base to the acid solution, the observed pH is 2.76. At this pH degree of protonation is

(1) 0%. ✗ (2) 10%. ✗

(3) 90%. (4) 99%. ✓

$pH = 2.76 < pKa = 4.76$

Protonated form will be more than 50%.

$pH = pKa - 2$

Solution :

$\text{Antilog}(pH - pKa) = \frac{DP}{P}$

$\text{Antilog}(2.76 - 4.76) = DP/P$

$\text{Antilog} -2 = DP/P$

$\frac{1}{100} = \frac{DP}{P}$

$DP : P = \text{Total}$

$1 + 100 = 101$

$P = \frac{100}{101} \times 100 = 99\%$

Buffers are independent of limited dilution

- 10 fold or 100 fold dilution of buffer with water will not change pH
- Because concertation of base form and acid form both get diluted

$$pH = \overset{5.0}{pK_a} + \log\frac{[A^-]}{[HA]}$$

$\frac{0.1\ M}{0.1\ M}$ $\xrightarrow[\text{Dilution}]{\text{100 fold}}$ $\frac{0.1/100}{0.1/100} = \frac{\cancel{0.001}}{\cancel{0.001}}$ 1

pH = 5 + log $\frac{\cancel{0.1}}{\cancel{0.1}}$

pH = 5

pH = pKa + log 1

pH = 5 + 0

pH = 5

Apply your mind

One litre of 0.2 M acetate buffer of pH=4.76 (pKa=4.76) was dilute 10 times, what will be altered pH?

(1) 3.76
(2) 4.76
(3) 5.76
(4) 6.76

pH = pKa

Base = Acid = Buffer

$\frac{0.1}{10}$ $\frac{0.1}{10}$ 0.2 M

0.01 0.01 = 0.02 M

$pH = pKa + \log \frac{B}{A}$

$= 4.76 + \log \frac{0.01}{0.01}$

$= 4.76 + \log 1$

$= 4.76$

Very high dilution affects pH of buffer

- Million fold or 10 Million fold dilution of buffer with water change pH
- Because concertation of base form and acid form in buffer become lesser than 10^{-7} M.
- Thus, H_3O^+ and OH^- of water will decide buffer
- It will closer to 7
    - For acidic buffer on high dilution it will be slightly lesser than 7
    - For basic buffer on high dilution it will be slightly higher than 7

Apply your mind

10 mM acetate buffer (pH 4.00) is diluted one million times with distilled water (pH 7.00). pH of this diluted buffer is:

(1) 4.00 ✗ (2) 7.04 ✗

(3) 8.00 ✗ (4) $\boxed{6.96}$ ✓

(Help: $\log_{10} 10^x = x$; $\log_{10} 1.10 = 0.04$; $\log_{10} 1.01 = 0.004$)

$10 \times 10^{-3}\,M = 10^{-2}\,M$

Dilution $\frac{10^{-2}}{10^{6}} = \boxed{10^{-8}}$

Buffer strength (Molarity) depends on moles/l of acid form and base form?

100 mL of 1 M Acetic acid + 100 mL of 1 M Acetate = pH of buffer = 4.76

100 mL of 1 mM Acetic acid + 100 mL of 1 mM Acetate = pH of buffer = 4.76

pH is same but strength is not same

strength of first buffer is more

$0.2 M = \frac{0.2 \text{ mole}}{1000 mL}$

$1000 mL \rightarrow 0.2$ mole

$1 ml \rightarrow \frac{0.2}{1000}$ mole

Apply your mind

A buffer is prepared by mixing 50 ml of 0.20 M acetic acid and 50 ml of 0.20 M sodium acetate. The pH of the buffer is 4.75 which is equal to the value of pKa of the acid. The molarity of the buffer is

$50 ml \rightarrow \frac{0.2}{1000} \times 50 = \frac{10}{1000} = 0.01$

(1) 0.10.
(2) 0.15.
(3) 0.20. ✓
(4) 0.40.

50 mL 0.2 M acetic acid
50 mL 0.2 M acetate

moles of
acetic acid = 0.01 in 50 mL
acetate = 0.01 in 50 mL

100 mL	Total = 0.02 mole in 100 mL
1000 mL	Total = 0.2 mole in 1000 mL

molarity. = 0.2 M = 0.2 mole/litre

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